Chemistry 129 Quiz #9

Name: Key

Solve the following problems:

1. Consider a 500.0mL buffer solution that is 0.160M NH_3 and 0.140M NH_4Cl.

$K_b = 1.8\times10^{-5}$ of NH_4^+

$$K_a = \frac{1.0\times10^{-14}}{1.8\times10^{-5}} = 5.6\times10^{-10}$$

(a) What's the pH of the buffer solution?

$$pH = pK_a + \log\frac{[\text{base}] \leftarrow NH_3}{[\text{acid}] \leftarrow NH_4^+}$$

$$= -\log(5.6\times10^{-10}) + \log\frac{(0.160)}{(0.140)} = \underline{9.31}$$

(b) Determine the pH of the solution after the addition of 0.020mol of HCl.

	NH_3	+ HCl	$\longrightarrow NH_4Cl$
BA	0.0800 mol		0.0700 mol
A		0.020 mol	
AA	0.0600 mol		0.0900 mol

$$pH = -\log(5.6\times10^{-10}) + \log\frac{(0.0600)}{(0.0900)} = \underline{9.08}$$

2. A 30.0mL sample of 0.230M cyanic acid (HCNO) solution is titrated with a 0.300M LiOH solution. Calculate the pH of the solution after the following volumes of base have been added: (a) 0.0mL, (b) 7.00mL, (c) 11.5mL, (d) at the equivalence point, (e) 24.0mL. Make a rough sketch of the titration curve including labels.

$K_a = 2\times10^{-4}$

$$V_e = V_b = \frac{C_aV_a}{C_b} = \frac{(30.0\,mL)(0.230)}{0.300\,M} = 23.0\,mL\ LiOH$$

(a) $V_b = 0\,mL$ LiOH (weak acid soln)

$$[H_3O^+] = \sqrt{K_a \times [HNCO]} = \sqrt{(2\times10^{-4})(0.230)} = 7\times10^{-3}\,M$$

$$pH = -\log(7\times10^{-3}) = \underline{2.2}$$

(b) $V_b = 7.00\,mL$ LiOH (Buffer Region)

	HNCO	+ LiOH	$\longrightarrow$ LiNCO	+ H_2O
BA	6.90 mmol			
A		2.10 mmol		
AA	4.80 mmol		2.10 mmol	

$$pH = pK_a + \log\frac{[\text{base}] \leftarrow NCO^-}{[\text{acid}] \leftarrow HNCO}$$

$$= -\log(2\times10^{-4}) + \log\frac{(2.10)}{(4.80)}$$

$$pH = \underline{3.3}$$

(c) $V_b = 11.5$ mL LiOH (Buffer Region)

$V_b = \frac{V_e}{2}$ ⇐ midpoint of titration   So pH = pKa

$pH = -\log(2\times10^{-4}) = \underline{3.7}$

(d) $V_b = 23.0$ mL LiOH (Equivalence Point − Weak Base Soln)

$HNCO + LiOH \longrightarrow LiNCO + H_2O$

	HNCO	LiOH	LiNCO	H_2O
BA	6.90 mmol			
A		6.90 mmol		
AA	0	0	6.90 mmol	

$K_b = \frac{1.0\times10^{-14}}{2\times10^{-4}} = 5\times10^{-11}$

$[NCO^-] = \frac{6.90 \text{ mmol}}{53.0 \text{ mL}} = 0.130 M$

$[OH^-] = \sqrt{K_b \times [NCO^-]} = \sqrt{(5\times10^{-11})(0.130)} = 3\times10^{-6} M$

$pOH = -\log(3\times10^{-6}) = 5.6$ $pH = 14.00 - 5.6 = \underline{8.4}$

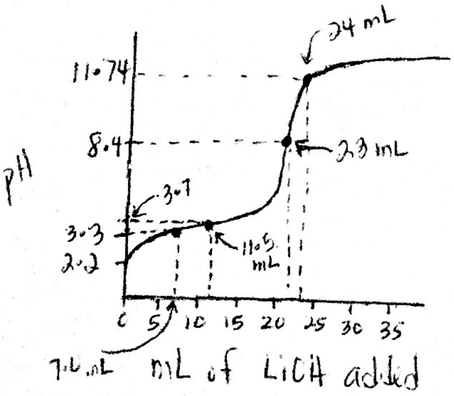
(e) $V_b = 24.0$ mL LiOH (Excess Strong base Soln)

$HNCO + LiOH \longrightarrow LiNCO + H_2O$

	HNCO	LiOH	LiNCO	H_2O
BA	6.90 mmol			
A		7.20 mmol		
AA		0.30 mmol	6.90 mmol	

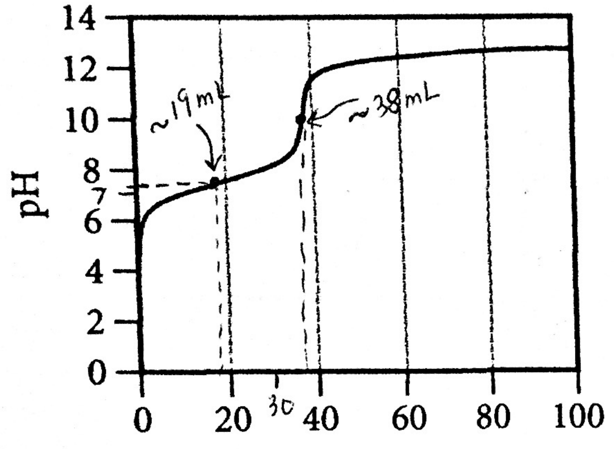
$[OH^-] = [LiOH] = \frac{0.30 \text{ mmol}}{54.0 \text{ mL}} = 0.0056 M$

$pOH = -\log(0.0056) = 2.26$ $pH = 14.00 - 2.26 = \underline{11.74}$

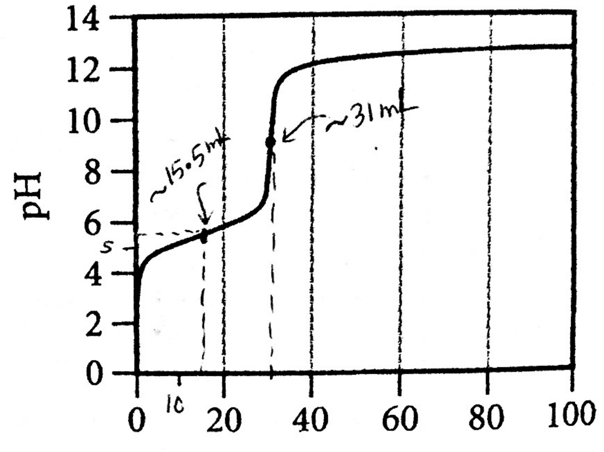


3. Consider the titration curves for two weak acids, both titrated with 0.100M NaOH: Oops!

Assume volume of acid is 30.0 mL



(a)



(b)

(i) Which acid solution is more concentrated?

(ii) Which acid has the larger Ka?

(i) Concentration of (a)

$C_aV_a = C_bV_b$

$$C_a = \frac{(0.100M)(38mL)}{30.0\ mL} = 0.13M$$

Concentration of (b)

$C_aV_a = C_bV_b$

$$C_a = \frac{(0.100M)(31mL)}{30.0\ mL} = 0.10M$$

* Acid (a) is more concentrated

(ii) Midpoint pH = pKa

Ka of (a)

At midpoint, pH ≅ 7.3

$$K_a = 10^{-pK_a} = 10^{-7.3} = 5 \times 10^{-8}$$

Ka of (b)

At midpoint ≅ pH = 5.6

$$K_a = 10^{-5.6} = 3 \times 10^{-6}$$